

# **TOPIC 3 – PERIODICITY**

## **3.2 – PHYSICAL PROPERTIES**

IB Chemistry  
T03D02



# 3.1 – Physical Properties – hrs

- 3.2.1 Define the terms first ionization energy and electronegativity. (1)
- 3.2.2 Describe and explain the trends in atomic radii, ionic radii, first ionization energies, electronegativities and melting points for the alkali metals (  $\text{Li} \rightarrow \text{Cs}$  ) and the halogens (  $\text{F} \rightarrow \text{I}$  ). (3)
- 3.2.3 Describe and explain the trends in atomic radii, ionic radii, first ionization energies and electronegativities for elements across period 3. (3)
- 3.2.4 Compare the relative electronegativity values of two or more elements based on their positions in the periodic table.



## 3.2 Electronegativity and Ionization E

### 3.2.1 Define the terms first ionization energy and electronegativity. (1)

- The **First Ionization Energy** is the minimum energy required to remove one mole of electrons from one mole of gaseous atoms (under STP conditions)
  - In general  $X(g) \rightarrow X^+(g) + e^-$
  - For example:
    - $H(g) \rightarrow H^+(g) + e^- \quad \Delta H = +1310 \text{ kJ mol}^{-1}$
    - It takes 1310 kJ of energy to strip away the first and outermost electron from one mole of hydrogen
  - The **Second Ionization Energy** would be the energy required to take the second electron and is very dependent on the placement in the PT.



# Electronegativity

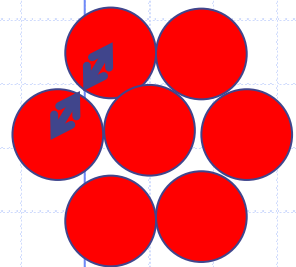
- The **Electronegativity** of an atom is the ability or power of an atom in a covalent bond to attract shared pairs of electrons to itself.
  - These values are based on the **Pauling scale** and range from 4.0 (F) down to 0.7 (Cs,Fr)
  - Values are pure numbers and simply relative to one another and do not hold units



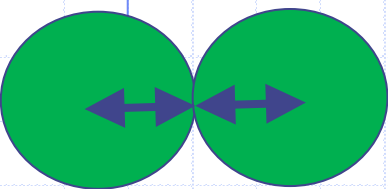
## 3.2

Trends Down Groups 1 & 7

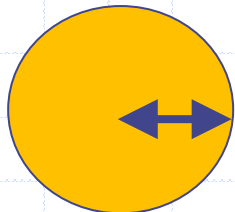
3.2.2 Describe and explain the trends in atomic radii, ionic radii, first ionization energies, electronegativities and melting points for the alkali metals ( Li  $\rightarrow$  Cs ) and the halogens ( F  $\rightarrow$  I ). (3)



Metallic radius



Covalent radius



Van der Waals radius



- For **metals**, the atomic radius is measured by half the distance between two atoms in a metal lattice
- For **non-metals in covalent** bonds, the atomic radius is measured by half the distance between the nuclei of two covalently bound atoms
- For **noble gases**, the atomic radius is the distance between the nuclei and the outermost electron of an isolated atom

# Effects on Atomic Radius

- Atomic Radius is determined by two opposing factors
  - The **Shielding Effect** by the electrons of the inner shell(s)
    - The shielding effect is the result of repulsion between the electrons in the inner shell and those in the outer or valence shell
  - The **Nuclear Charge**, the number of protons
    - The **Effective Nuclear Charge** is the relative number of protons in the nucleus to electrons it must interact with. The larger the ratio of protons to electrons, the more effective the nucleus will be and the relative result will be a smaller atomic radius



# Effects on Atomic Radius (2)

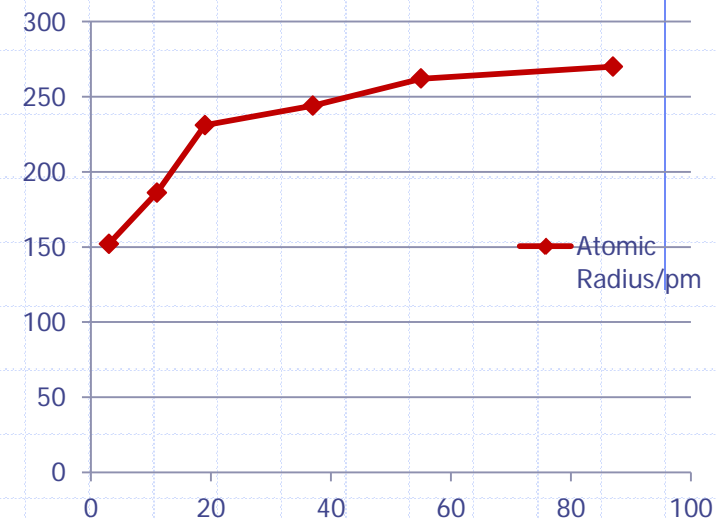
- When moving down the periodic table, the radius increases due to:
  - An increase in the number of complete electron shells between the nucleus and outer electrons
  - An increase in the shielding effect of the outer electrons by the inner electrons



## 3.2

# The variation of atomic radii in the **Alkali Metals**

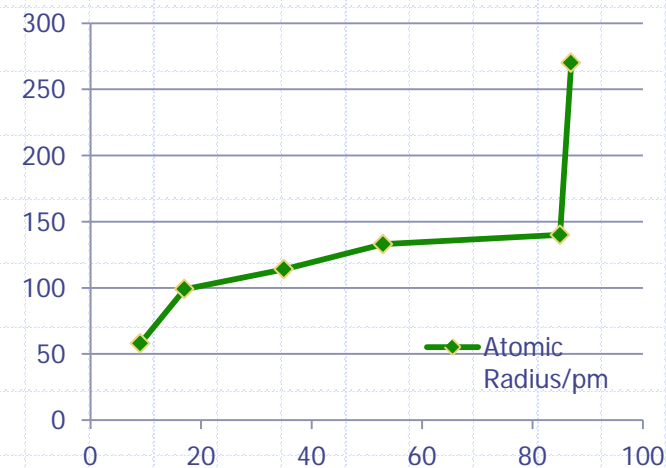
Atom	Atomic Number	Atomic Radius/pm
Li	3	152
Na	11	186
K	19	231
Rb	37	244
Cs	55	262
Fr	87	270



# The variation of atomic radii in the **Halogens**

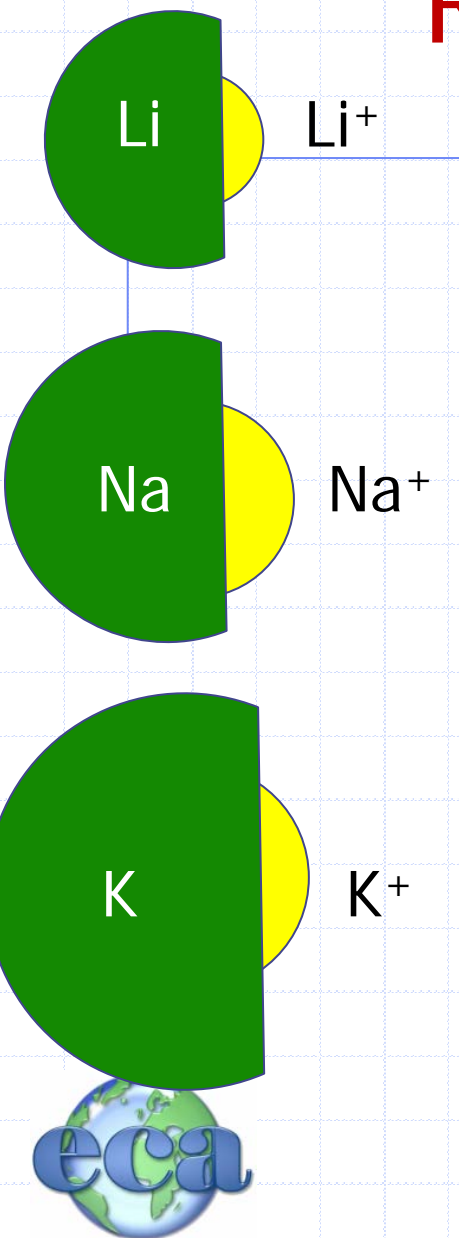
Atom	Atomic Number	Atomic Radius/pm
F	9	58
Cl	17	99
Br	35	114
I	53	133
At	85	140
Fr	87	270

Atomic Radius/pm

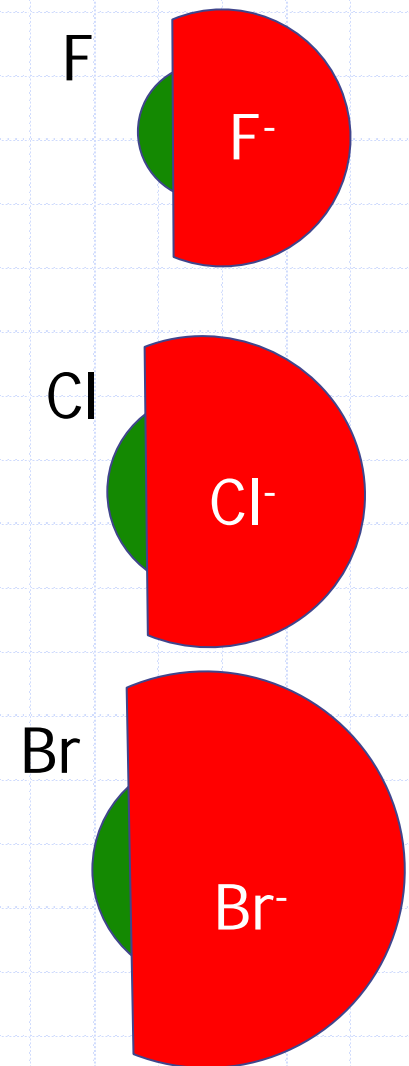




# Radii of ionic species



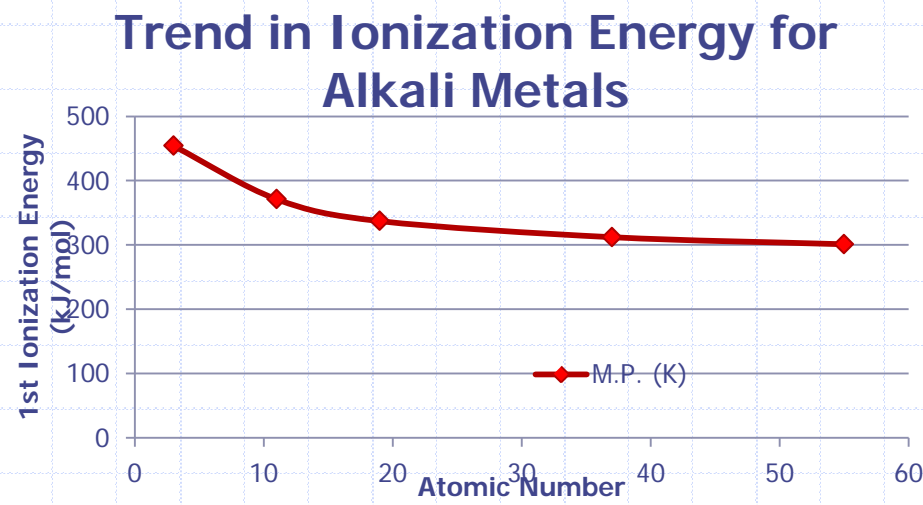
- Radius increases down a group
- Radius decreases across a period
- Cations are smaller as they have an increase effective nuclear charge
- Anions are bigger as they have a decreased effective nuclear charge



# Trend in 1<sup>st</sup> ionization energy

- Down a group, as atomic radius increases, and increased electron shielding is present, the valence electrons are “forgotten”
  - Further out from the nucleus
    - Less attraction between protons and valence electrons
    - More electron shielding (repulsions)
  - Requires less energy to remove an electron

Atom	Atomic Number	1 <sup>st</sup> I.E. kJ mol <sup>-1</sup>
Li	3	519
Na	11	494
K	19	418
Rb	37	402
Cs	55	376



# Trend in Electronegativity

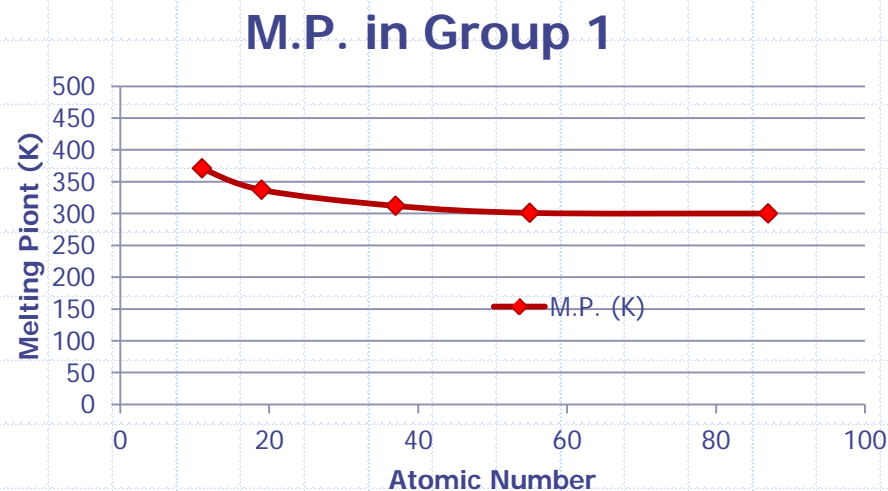
- Values decrease down a group
  - Due to an increased radius and distance between the nucleus and shared pairs of electrons.
  - Nuclear charge is increased but is counteracted by additional electron shielding
- Values increase across a period
  - As more electrons are added to the valence shell, the nucleus has more of an attraction, the radii decreases, and stability is nearer.



# Melting Point in Alkali Metals

- The M.P. of group 1 decreases down the group.
  - Held together by metallic bonds and a lattice of positive ions with a sea of delocalized electrons
  - The M.P. decreases down the group because
    - The metallic bond strength decreases
    - Attractive force between delocalized electrons and nucleus decreases due to the atomic radius increase

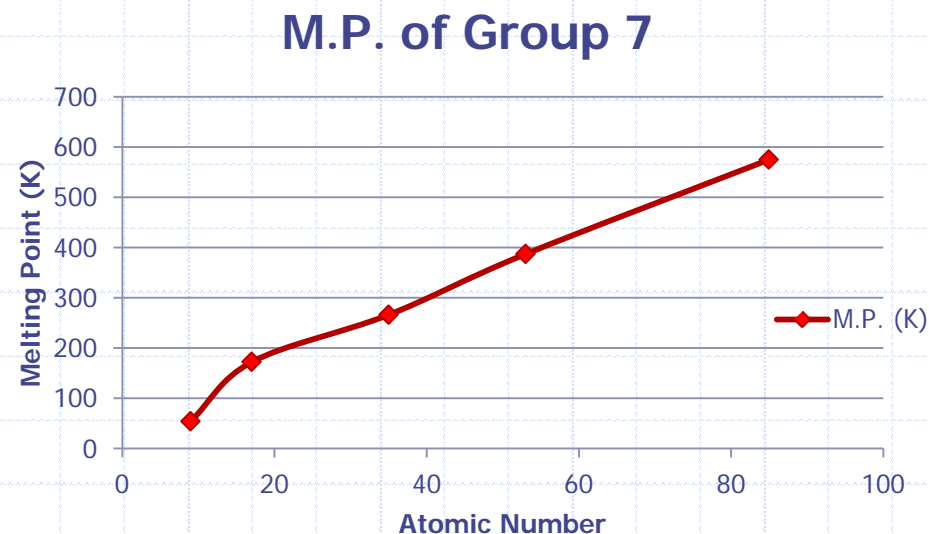
Atom	At. #	M.P. (K)
Li	3	454
Na	11	371
K	19	337
Rb	37	312
Cs	55	301
Fr	87	300



# Melting Point in Halogens

- The M.P. of group 7 increases down a group
  - This is because molecules become large and the attractive forces (Van der Waals forces) increase between them.
  - Increase with mass and number of electrons

Atom	At. #	M.P. (K)
F	9	54
Cl	17	172
Br	35	266
I	53	387
At	85	575



## 3.2 Trend Across 3<sup>rd</sup> Period

3.2.3 Describe and explain the trends in atomic radii, ionic radii, first ionization energies and electronegativities for elements across period 3. (3)

- There is a gradual decrease in **atomic radii** across period 3.
- As one more electron and proton are added across:
  - Electrons are added to the same shell and therefore only a slight increase in  $e^-$  shielding
  - Additional protons in the nucleus increases the nuclear charge and the electrons in the valence shell are pulled more closely



# Ionic Radii Across Period 3

- The radii of positive ions decrease from  $\text{Na}^+$  to  $\text{Al}^{3+}$
- The radii of negative ions decrease from  $\text{P}^{3-}$  to  $\text{Cl}^-$
- The ionic radii increase from the  $\text{Al}^{3+}$  to  $\text{P}^{3-}$

Element	Na	Mg	Al	Si	P	S	Cl
Ion	$\text{Na}^+$	$\text{Mg}^{2+}$	$\text{Al}^{3+}$	$(\text{Si}^{4+}, \text{Si}^{4-})$	$\text{P}^{3-}$	$\text{S}^{2-}$	$\text{Cl}^-$
Ion Radii (pm)	98	65	45	(42, 271)	212	190	181

decreases

decreases

Large increase as  
direction of charge  
changes

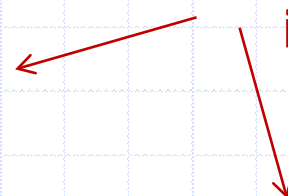


# Isoelectronic Species

- **Isoelectronic Species** are atoms that have the same number of electrons.
- The **Effective Nuclear Charge** is the ratio protons to electrons. The greater the ratio, the smaller radius will be relative to it's neutral state

Species	Na <sup>+</sup>	Mg <sup>2+</sup>	Al <sup>3+</sup>
Nuc. Charge	+11	+12	+13
# electrons	10	10	10
Ion Radii (pm)	98	65	45

Each are individual sets (of 3)  
isoelectronic species



Species	P <sup>3-</sup>	S <sup>2-</sup>	Cl <sup>-</sup>
Nuc. Charge	+15	+16	+17
# electrons	18	18	18
Ion Radii (pm)	212	190	181

The two sets cannot be compared as the second set has another complete shell





# 1<sup>st</sup> Ionization Energy Across Group 3

- There is an increase in ionization energy across period 3
  - Due to increased nuclear charge in comparison to a minor increase in the shielding effect
  - The shells are pulled more closely together and the nucleus has a greater attraction to the valence electrons
    - This increases the ionization energy
    - Some decreases can be explained by viewing the electron configuration

Element	Na	Mg	Al	Si	P	S	Cl
1 <sup>st</sup> IE kJ mol <sup>-1</sup>	494	736	577	786	1060	1000	1260

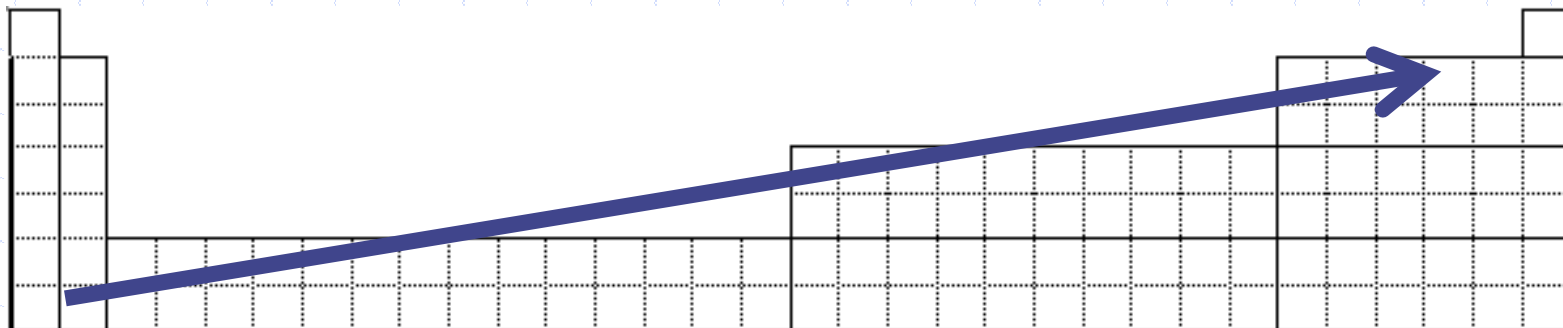
## 3.2

# Comparison of Electronegativities

3.2.4 Compare the relative electronegativity values of two or more elements based on their positions in the periodic table.

- As discussed, electronegativity increases across a period and decreases down a group.
- This can be explained just as 1<sup>st</sup> ionization energy
  - Increased nuclear charge compared to small increase in the shielding effect

Element	Na	Mg	Al	Si	P	S	Cl
Electronegativity	0.9	1.2	1.5	1.8	2.1	2.5	3.0



# Comparison of two Elements

- A comparison of electronegativities can help to determine the type of bonding that will occur
  - Far away from one another on the P.T.
    - Large difference in electronegativities
    - If a metal and non metal = ionic bond
    - Metal and a metal = metallic bond
  - Close to one another on the P.T.
    - Small differences in electronegativities
    - Polar, and possibly non-polar covalent bonding

